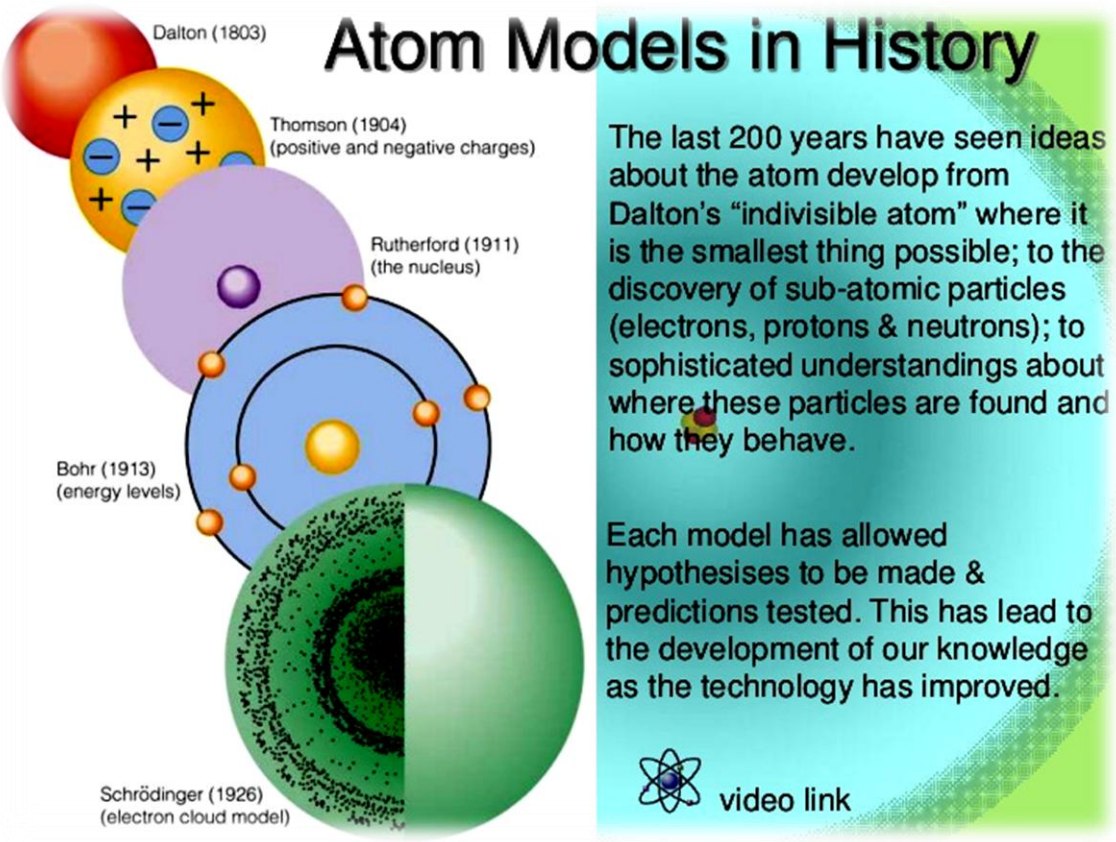
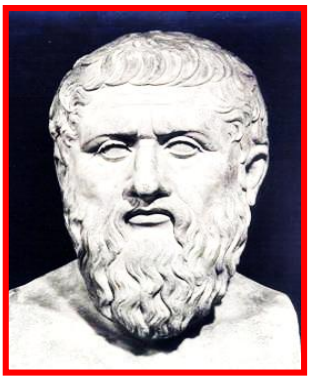
Chapter one

Atomic structure



In the following we illustrate the historical evolution for atomic structure concept:

1-Democrits's idea (Greek philosophers)



Any piece of matter can be divided into smaller parts and each part can be subdivided into smaller parts which can't be divided this part is called Atom.

(ATOM in the Greek language A = no , TOM= divide)

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2-Aristotle

a- He refused the idea of Greek philosophers about the atom.

b- He supposed that all matters composed of 4 constituents which are (water, air , dust, and fire

c- He postulated that the cheap metals such as iron or copper can be changed into precious ones like gold by changing the percentage of four constituents.

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3-Boyl (Irish 1661)

I-He refused the Aristotle concept.

Ii-He was the 1st scientist to define the element

The element:

A pure simple substance which cannot be changed into simpler form by the traditional chemical methods.

4- Dalton' s atom (English 1803)

He supposed that:

1) The element is composed of very minute particles which are named atoms.
2) Each element is a solid undividable atoms.
3) Atoms of the same element are similar in mass but differ from atoms of other elements.
4) Compounds are formed by the combination of atoms of different elements in a simple numerical ratio.

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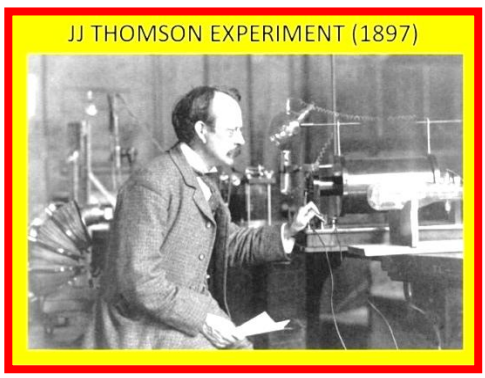
5- Thomson's atom 1897



Thomson discovered the cathode rays

Cathode- rays experiment

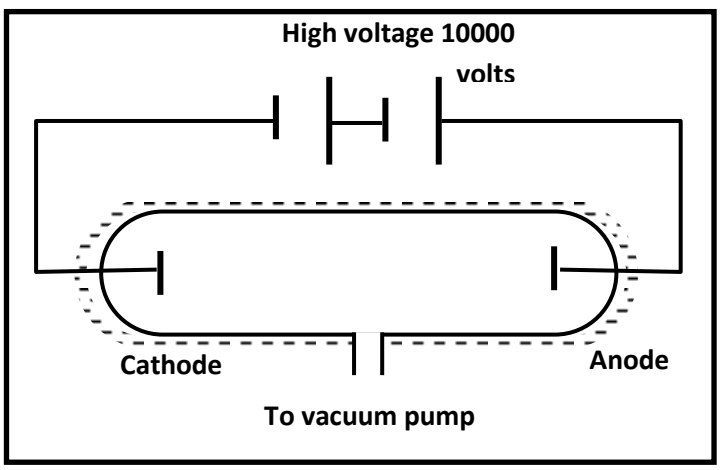
(Discovery of the electron)

a) All gases under normal conditions of pressure and temp don't conduct electricity.

b) If a glass tube evacuated from the gas to a very low pressure

c) The gas will conduct electric current.

d) If the potential difference between the two poles increases up 10000 Volts



e) A Flow of invisible rays are emitted from the cathode causing glowing to the wall of the tube and called cathode rays.

Cathode rays

A stream of invisible rays are emitted from the cathode causing glowing to the wall of the tube when the potential difference between the two poles increases up 10000 volts and under very low pressure

Properties of cathode rays

1) Consist of tiny particles have mass and velocity.

2) Transfer in straight lines glowing the glass facing the cathode.

3) Have negative charge.

4) Have a thermal effect.

5) Affected by electric and Magnetic field.

6) Take part in the structure of all substances.(G.R)

Bec. Cathode rays don t change in behavior by changing either cathode material or type of the gas

Thomson's model

He conclude from the last experiment that

i- The atom is a homogenous sphere of positive electricity.

ii- Inside it there are negative electrons enough to make it electrically neutral.

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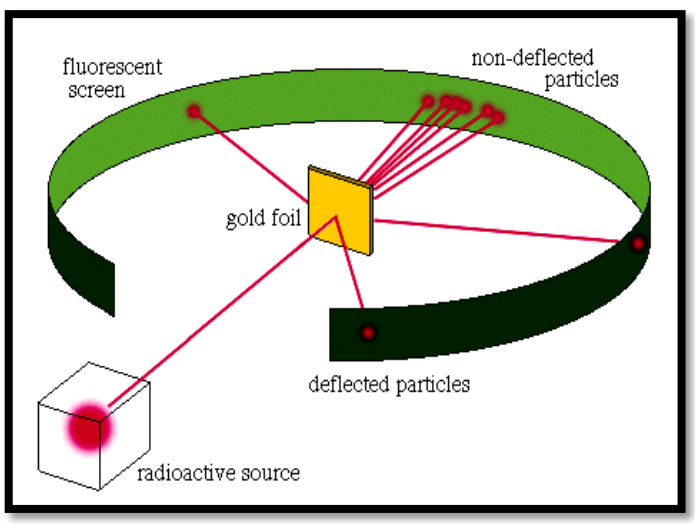
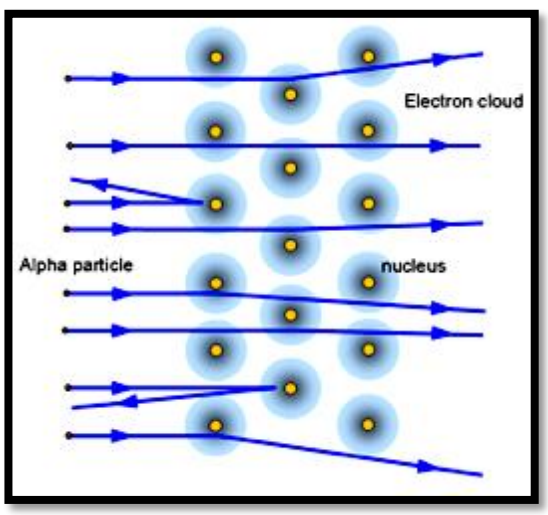
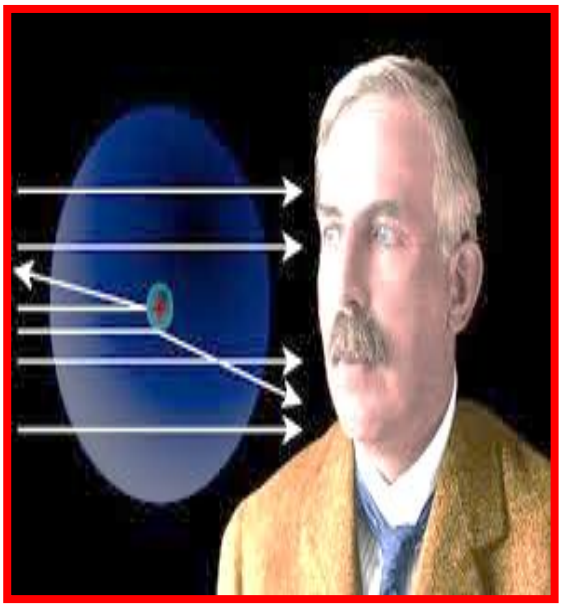
6-Ruther ford's model of the atom





6-Ruther ford's experiment

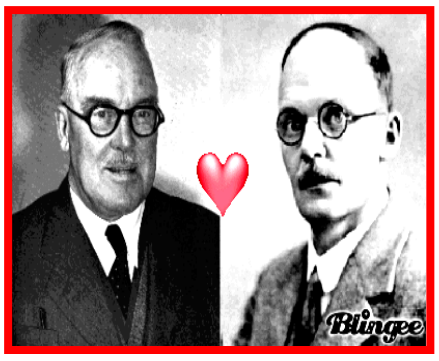


1 – He allowed alpha particles to hit a metallic plate lined with Zinc sulphide **(glows when hits with alpha rays)**

2 – On placing a gold foil in the front of alpha rays he concluded the following from the following observation.

Observation	Result
1- Most of alpha particles appeared in the same position before putting the gold foil.	1- Most of the atom is a space not solid as explained by Dalton and Thomson.
2-A very small percentage of alpha particles reflected back to appear as flashes on the front of the foil	2- The atom has very small part with very small volume but high density. (was named the nucleus)
3- Some flashes appeared on the sides of 1st site.(deflected)	3- The dense parts of the atom which concentrate in it most mass have same charge of alpha particle (+ve).

Note



In 1911 **Geiger and Marsden** (Rutherford's student) performed a famous experiment According to suggestion of Ruther ford by the following apparatus.

Rutherford designed his atomic as follow:-

1- The Atom:

a) Although it has very small size but it has a complicated structure that resembles the solar system

b) In which electrons revolve around the central nucleus in orbits as planets revolve around the sun.

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2- The Nucleus:

It is much smaller than the atom. Located in the centre of the atom with (+ve) charge.

(G.R)

……..……

There is a big space between the nucleus and orbits of electrons

So most of the atom is a space.

3- Electrons:

1- Have negligible mass compared to that of the nucleus.

2- No. of electrons (-ve) equals no of protons (+ve)

So the atom is electrically neutral.

3- Electrons revolve around the nucleus in a fixed orbit as electrons are affected by two forces equal in strength but in opposite direction, which are :

a- Force of attraction of the nucleus to electrons.

b- Centrifugal force due to velocity of electron around the nucleus.

Give reason

Electrons are not attracted to the nucleus (don't fall into nucleus)?

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Most mass of the atom is concentrated in the nucleus

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Sheet 1

A) Write the scientific expression for each statements of the following :

1- A pure substance that can't be divided into a simplest one using chemical methods.

2- The atom is a homogeneous positively charged ball and electrons are embedded in it.

3- A stream of invisible rays that affect the wall of electric discharge tube.

4- Particles that cause flashing when falling on a metallic plate lined with a layer of zinc sulphide.

5- A stream of invisible rays which are emitted in discharge tube under very low pressure and a potential difference of about 10000 volts.

B) Mention the name of the scientist who:

1- Believed that all substances are composed of water, air, dust and fires.

2- Confirmed that the element is a pure substance which can't be divided into a simple one.

3- Assumed that the element is composed of similar tiny dense atoms.

4- Imagined the possibility of dividing any piece of matter to smaller undividable fragment , he named it an "atom".

C)Choose the correct answer for each of the following sentences:

1- All the following are from the properties of the cathode rays except they ……..

a) have a thermal effect b) move in straight lines
c) are positively charged
d) are effected by both electric and magnetic fields

2- Geiger and ……. performed Rutherford's experiment.

a) Marsden b) Dalton c) Rutherford d) Thomson

3- From the properties of the cathode rays that ……

a) have mass only b) have charge only

c) have no mass and charge d) have mass and charge

4- ……. rays are deflected towards the positive pole when it is affected by the electric field.

a) Alpha b) The cathode c) Gamma d) X

5- **The first scientist who defined the element is.....**

a. Boyle b. Rutherford c. Bohr d. Thomson.

6- **All matters are composed of four components (water, air, dust, and fire) with a different ratio. That idea belongs to .....**

a. Bohr b. Rutherford c. Dalton d. Aristotle.

7- **The strong evidence that proved that cathode rays exist in all matters .....** a. they have thermal effect.

b. flow in straight lines.

c. consist of fine particles.

d. they have the same properties and behavior whatever the gas or the cathode material used.

8Cathode rays consists of particles called .....

a. alpha particles b- electrons c - atoms d – orbitals

<u>Give reason for each of the following:</u>

1- The old thought that iron can be converted into gold.

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2- Alpha particles deviate in the opposite direction of cathode ray deviation , when they are exposed to electric or magnetic fields.

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3- Zinc sulphide is used to detect invisible alpha particles.

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4- The mass of the atom is concentrated in the nucleus.

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5- The atom is electrically neutral.

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Explain the observations upon which Rutherford reached the following conclusions:

a. Most of the atom is an empty space and it is not a solid sphere.

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b. There is a very dense tiny piece of the atom later on named the nucleus.

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c. The charge of the dense part of the atom in which most of its mass is concentration should have a positive charge similar to alpha particles.

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Explain Thomson's atomic model.

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Atomic emission spectra

Studying and explaining the atomic spectrum was the key that solved the puzzle of the atomic structure

In 1913by (Niels Boher) and deserved noble prize in 1922. Atomic emission spectrum:

1 – By heating gases or vapours of substances to a high temperature (by heat or electric spark) under low pressure it produces light.

2 – By using **spectroscope** we find that this light consists of a fixed number of colored lines called line spectra.

Line spectra:

A type of spectrum composed of small number of restricted colored lines separated by dark areas

Line spectrum of hydrogen atom:

It appears as four colored lines separated by dark areas

Give reason:

The emission spectrum is named as line spectrum?

Bec.it is composed of restricted no. of colored lines which are separated by dark areas.

The spectral lines are essential characteristics for each element?

Bec. there is no two elements have the same spectral lines it looks like the finger print

Note:

Line spectrum of sun rays shows that composed of hydrogen and Helium.

7- Bohr's atomic model

A- <u>Points that agree with Rutherford's postulates</u>

1- A positively charged nucleus exists in the center of the atom.

2- The number of negative electrons equal the number of positive protons in the nucleus.(protons)

3- Electrons revolve around the nucleus in orbits due to centrifugal and attraction forces.

B- <u>Bohr's postulates (New postulates):</u>

4- Electrons orbit □ the nucleus in a rapid movement without gaining or losing energy (stable atom).

5-Electrons orbit□ the nucleus only in definite allowed energy levels, so they can't be found at intermediate distances.

6- Each electron in the atom has a definite amount of energy depending on the distance between its E.L and nucleus.
(This energy increases as its radius increases.)

7- It was found that the maximum no of energy levels in the heaviest known atoms in their ground state (unexcited) is only seven (K, L, M, N, O, P, Q). Each level has energy expressed by a whole no called principle Q. No.

<u>Ex:</u> The 1st E. Level K its principle Q. no = 1

The 2nd E. level L its principle Q. no = 2

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<u>Notes:</u>

1- The difference in energy between energy levels not equal

2- So the difference in this energy decreases further from the nucleus.

3- This means that the quantum of energy required to transfer an electron from one energy level to another is not equal.(G.R)

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4- The electron does not move from its level to another unless the energy absorbed or emitted is equal to the difference in energy between 2 levels i.e. one quantum.
(There is no half quantum for instance). Can't be divided or doubled

Give reason:

It is wrong to say that e' to be transferred from E.L (K) to E.L (M) needs amount of energy equals 2 quantum.

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5- The energy levels in Bohr's atom may be represented by the steps of the stair (a ball is unstable in mid air between the steps . its only permitting to jump on the step . it is the same case in the electron transition between energy levels)

Excited Atom:

It is an atom that acquired an amount of energy by heating or an electric discharge

Advantages of Bohr:

1- It explained hydrogen atom spectrum.

2- He introduced the idea of quantized energy to detect energy of electrons in energy levels.

Disadvantages (inadequacy) of Bohr's atomic model:

1- He failed to explain the spectrum of any other element even that of Helium except hydrogen (Simplest Electronic System).

2- He considered the electron as a (-ve) particle only ,**and ignored its wave properties**

3- He postulated that it is possible to determine precisely both speed and location of an electron at the same time. **This is experimentally impossible**

4- He described the electron when moving in a circular planer orbit, which means that hydrogen atom is planer. **In fact hydrogen atom has a spherical shap (3 dimensional coordinates).**

8-The principle of modern atomic theory (modification of Bohr's model)

1 – Dual nature of the electron.

2 – The Heisenberg uncertainty principle.

3) Wave mechanical theory of the atom.

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1 – The dual nature of the electron

All the previously mentioned theories considered the electron just as a minute negatively charged particle. However, the experimental data showed that:
The electron has a dual nature **G.R**
Bec. It is a material particle which also has wave properties.

The dual nature of electron

The is a material particle which has wave properties

2- The Heisenberg uncertainty principle: (quantum mechanics)

It is practically impossible to determine both position and the velocity of the electron exactly (precisely) at the same time. We can only say that it is probably to a greater or lesser extent to locate the electron in this or in that place. This is to speak in terms of probability seems to be more precise.

Heisenberg uncertainty principle

The determination of both the velocity and the position of an electron at the same time is practically impossible and subjected to the laws of probability

3. The wave mechanical theory of the atom

(Schrodinger wave mechanics theory):
He applied the ideas of Planck, Einestein, De Broglie and Heisenberg so
On dissolving Schrodinger's equation it's possible to :

1- determine the allowed energy levels

Define the region of space around the nucleus where it is most probable to find the electron in each energy level.

Electron Cloud: (used to describe any orbital)

"Area of space around the nucleus where there is a great probability for finding electrons in all direction and all positions."

The difference between the orbit and orbital concepts according to both Bohr and the wave mechanics theories:

Electron cloud:
The region of space around the nucleus, in which the electron probable exists in all directions and distances (dimensions)

Orbital

The region within the electron cloud of high probability of finding the electron

On dissolving the mathematical solution of the Schrodinger equation introduced four numbers which are called quantum numbers.

Sheet 2

Choose the correct answer for each of the following sentences:

1- …… succeed to explain the line spectra which solved the puzzle of the atomic structure.

a) Heisenberg b) Bohr c) Kosel d) Haber

2- The study of hydrogen atomic spectra is considered the key that Bohr used to know ……..

a) that electrons are negatively charged

b) that atom contains a central nucleus

c) the energy levels of the atom d) all the previous

3- When an electron absorbs half quantum of energy , it will ……

a) transfer to a higher energy level b) transfer to a lower energy level

c) be in the same energy level d) no correct answer

4- When an electron transfers from an energy – level near to the nucleus to a further one , it will…….

a) lose a quantum b) gain a quantum

c) radiate energy d) don't lose a part of its energy

5- The energy required to transfer an electron from the first energy level to the second one is …….. that needed to transfer it from the second energy level to the third one.

a) equal to b) less than c) greater than

Give reason for each of the following:

1- Spectrum of hydrogen atom is composed of a group of spectral lines.

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2- The energies required to transfer an electron between the energy levels are not equal.

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3- the electron has a dual nature

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Choose the correct answer:

1-Heating gases or vapours to high temperature under reduced pressure, they.....

b. emit light a. absorb energy

d. emit Alpha radiation c. emit gamma radiation

2- - When an electron absorb a quantum of energy it .....

a. Transfers to all higher energy levels

b. Transfers to the higher energy level corresponding to the absorbed energy

c. Transfers to lower energy level

d. Transfers to the lower energy level that corresponds to the absorbed

3-The study of hydrogen atomic spectra is considered the key that Bohr used to know ……..

a.that electrons are negatively charged

b.that atom contains a central nucleus

c.the energy levels of the atom

d.all the previous

Write the scientific expression for each statement of the following:

1) It represents the number of orbitals within a certain energy sublevel and their direction in space.
2) It indicates the number of energy sub levels in each principle level.
3) Bohr used it in hydrogen atom it is used to define Orders of principle energy levels
4) They are numbers define the volume of space (orbital) where there is maximum probability of finding electrons. Besides, they define the energy, shape and direction of orbitals.
5) The region within the electron cloud of high probability of finding the electron
6) The region of space around the nucleus, in which the electron probable exists in all directions and distances (dimensions)
7) It is practically impossible to determine both position and the velocity of the electron exactly (precisely) at the same time.
8) It is an atom that acquired an amount of energy by heating or an electric discharge
9) A type of spectrum composed of small number of restricted colored lines separated by dark areas

Explain the contributions of the following scientists to chemistry :

a. Bohr

b. Heisenberg

c. Schrodinger

What is meant by

a. Electron cloud

b. The dual nature of electron

c. The Heisenberg uncertainty principle

Quantum numbers:

Quantum numbers:

They are numbers define the volume of space (orbital) where there is maximum probability of finding electrons. Besides, they define the energy, shape and direction of orbitals.

There are four quantum numbers:

1-The principle quantum number (n)
2-The subsidiary quantum number (ℓ)

3-The magnetic quantum number (mℓ)

4- The spin quantum number (ms)

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Principle quantum number (n)

a) Bohr use it in hydrogen atom it is used to define

1-Orders of principle energy levels and their number are seven in the heaviest one in its ground state.

-1st E.L	K	Is filled with	2 electrons
-2nd E.L	L	Is filled with	8 electrons
-3rd E.L	M	Is filled with	18 electrons
-4th E.L	N	Is filled with	32 electrons

2-The number of electron required to fill its level ($2n^2$)

This rule is not applied in levels higher than fourth level
Because the atom becomes unstable if the number of electrons exceed 32 electrons in any level.
* The principle quantum number must be a whole number exclude zero
*Each energy level is divided into sublevels.

Subsidiary quantum number (ℓ):

It indicates the number of energy sub levels in each principle level.

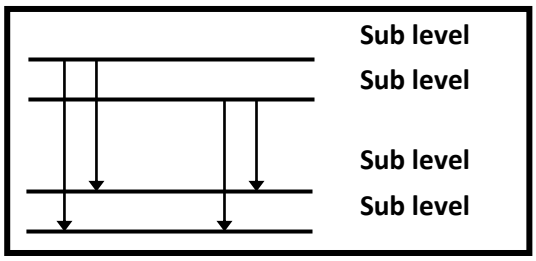
b)Each principle energy level contains a number of energy sublevels equal to its principle quantum number.

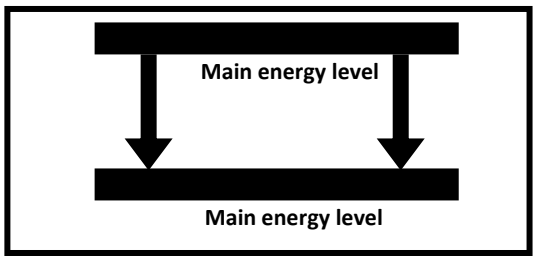
c) The energy sublevels take the symbols and values which are shown in the following table:

Symbols of sublevels	s	p	d	F
Values of subsidiary quantum number (l) [o: n-1]	0	1	2	3

d)They was noticed by **Sommerfield** that each spectral line is a number of fine spectral lines that represent electron transition between very near energy levels (sublevels).

e)There is a small difference in energy of sublevels s<p<d<f





Example1:

What are the probable (ℓ) values when n = 3

Solution

∵ n= 3 so the no. of sublevels = 3

So the probable (ℓ) values ranges between $[0 : (n-1)] = [0:(3-1)] = 0,1,2$

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Example2:

Mention the sublevels that exist in an atom knowing that its last principle level is L

Solution

P.Q.No.	**N**	**(ℓ)**	**Sublevels**
K	1	0	1S
L	2	0,1	2S,2P

So sublevels in this atom are (1S, 2S, 2P)

3-The magnetic quantum number

The magnetic number is characterized by the following:

a) It represents the number of orbitals within a certain energy sublevel and their direction in space.

b) It is represented by odd and integer numbers between (-ℓ,...,0,...+ℓ).

The following table explain probability of magnetic quantum number for atom (n=4).

(n)	(ℓ)	(m*l*)
1	0	0
2	0	0
	1	-1, 0, +1
3	0	0
	1	-1,0,+1
	2	-2,-1,0,+1,+2
4	0	0
	2	-2, -1, 0, +1, +2
	3	-3, -2, -1, 0, +1, +2, +3

c) Sublevel (s) has one orbital of spherical symmetrical shape around the nucleus.

The sublevel (p) consists of three orbitals whose axes take the three spatial orientations (orientation in space x, y, z). Thus they are designated as px, py and pz

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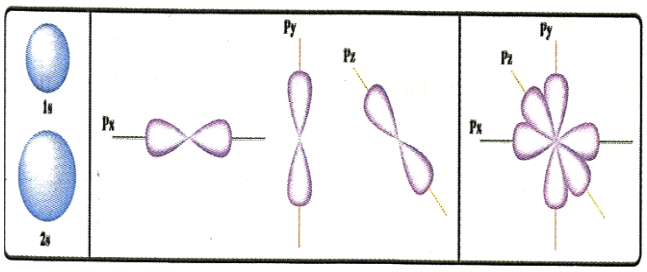
Note:

Each p orbital is perpendicular to the other two. The electron cloud of each orbital takes the form of two pears meeting head to head (dumb-bell shaped) at a node point of zero electron density.



sublevel (d) has 5 orbitals and sublevel (f) has 7 orbitals.

Give reason:

The sublevel P is completely filled with 6 electrons while the sublevel d is completely filled with 10 electrons?

Bec. The sublevel **P** contains 3 orbitals while the sublevel **d** contains 5 orbitals And each orbital filled with 2 electrons

Example1:

What are the probable (mℓ) values when ℓ = 2

(mℓ) values ranges between (-ℓ, 0 , +ℓ)

(mℓ) values ranges between (-2,-1, 0 ,1,2)

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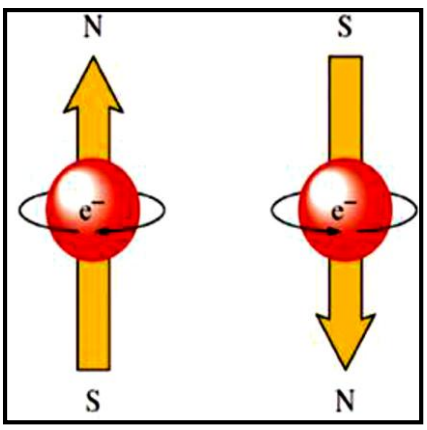
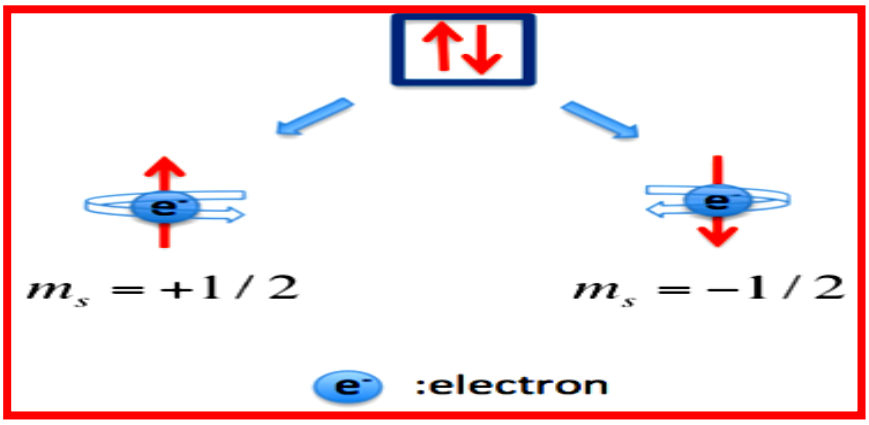
4-Spin quantum number (m_s)



- Any orbital cannot be occupied by more than two electrons.
- Each electron spins on its own axis during its orbit around the nucleus.
- This can be illustrated when we imagine the spinning of the earth on its own axis during its rotation around the sun.
- Although the two electrons of the same orbital carry the same negative charge, we might expect them to repel.
- Yet as a result of the spinning of each electron on its own axis, a magnetic field arises in a direction opposing the direction of the other magnetic field arising from the spinning of the other electron.
- It is said that the two electrons are in a spin paired state and these are designated as (↑↓)
- The following considerations must be observed about the spin quantum number:

1- It defines the type of spin motion of the electron and since the electron can only spin in either of the two directions i.e. clock-wise (↑) m_s = + 1/2 or anticlock-wise (↓) m_s = -1/2



Summary of the relationship between the principle energy level, sublevels and orbitals

1. The number of energy sublevels equals the number of the principal level to which they belong,

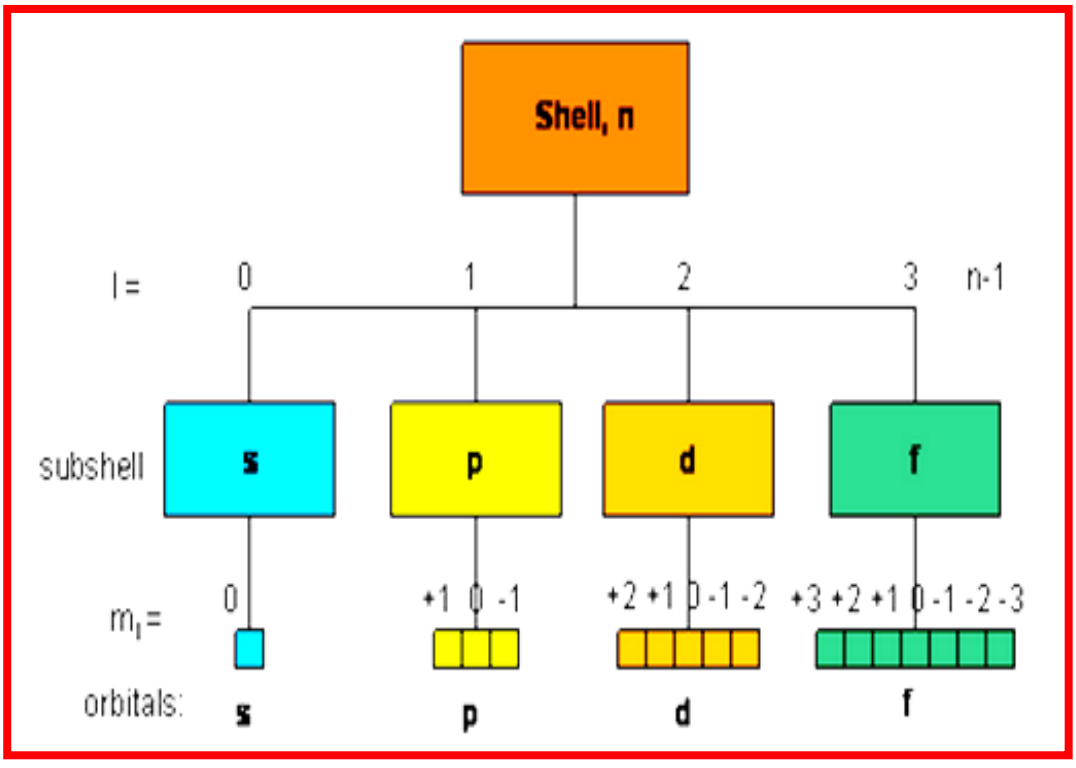
i.e. the first principal level consists of one sublevel and the 2^{nd} principal level has two sublevels ...etc.

2. The number of orbitals within a principal energy level square the number of level (n^2)

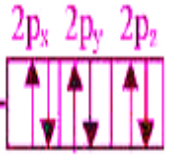
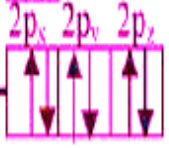
i.e. The 2^{nd} energy level consists of 4 orbitals 2s, $2p_x$, $2p_y$, $2p_z$ and the 3^{rd} energy level consists of nine orbitals (3S, 3px, 3py, 3pz and five 3 d orbitals).

3. The number of electrons occupying a given principal energy level equals two times the square of this level ($2n^2$).

For example, the 2^{nd} level can take eight electrons distributed as follows 2s2 , $2p^2_x$, $2p_y^2$, $2p_z^2$.



The quantum numbers of the electrons occupying the first three energy levels may be summarized as follows:

Level	Principal quantum no.(n)	Subsidiary quantum no.(ℓ)	Magnetic q. no. $m\ell = 2\ell + 1$
	Define the energy levels	Define the no. of sublevels, which equal the principal quantum no.(n)	Define the no. of orbitals in each sublevel. S=1 ,P=3 d=5 ,f=7
K	1	1S	
L	2	2S	
		2p	
M	3	3S	
		3P	
		3d	

Principles of distributing electrons

1. Pauli's exclusion principle: it states that:

It is impossible for two electrons in the same atom to have the same four quantum numbers.

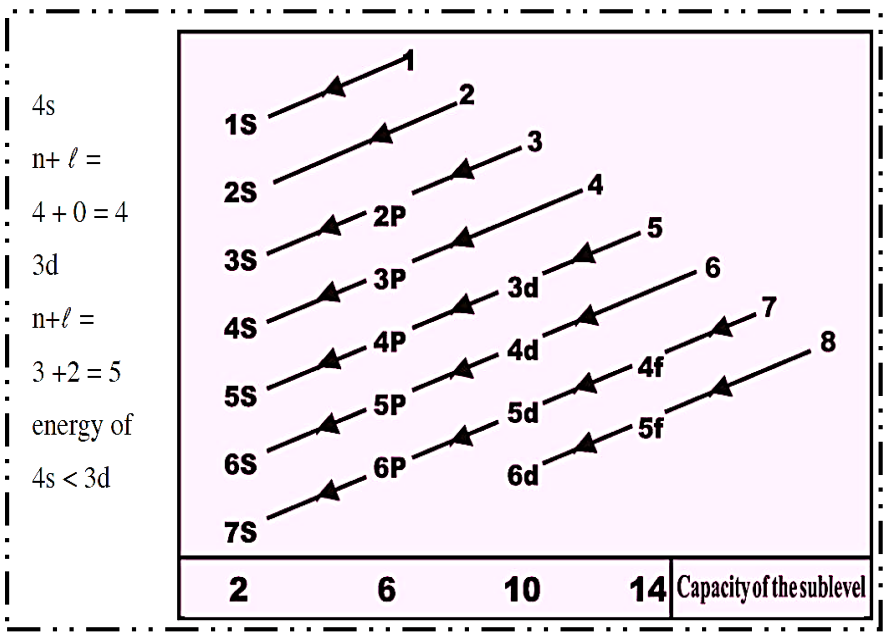
The following table explains two electrons of 3s similar in quantum numbers (n, ℓ, m_ℓ) but differ in (ms):

4 quantum numbers	n	ℓ	mℓ	m_s
first electron	3	0	0	$+\frac{1}{2}$
second electron	3	0	0	$-\frac{1}{2}$

2-Aufbau (building up) principle

It states that "Electrons occupied the energy sublevels in order of increasing their energy, the lowest sublevels are filled first"

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p



Examples:

$_{11}Na$	
$_{20}Ca$	
$_{30}Zn$	
$_{35}Br$	

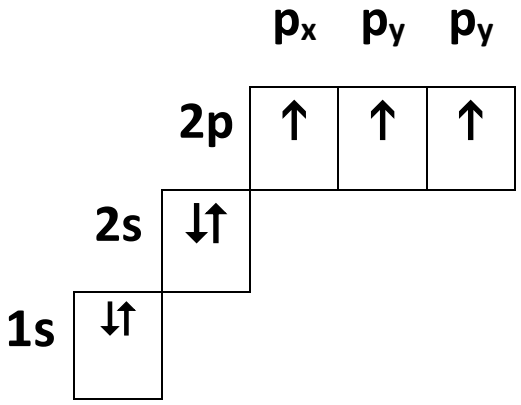
.....

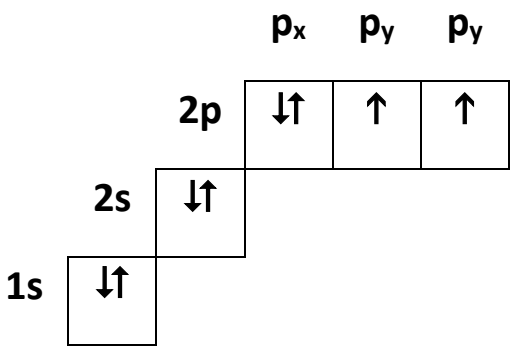
3-Hund's rule

States that "No electron pairing takes place in a given sublevels until each orbit contains one electron".

.....

Write the electronic configuration of $_{7}N$ and $_{8}O$ according to Hund's rule





.....

Give reasons:

1- **The sublevels 4S is filled by electrons before the sublevel 3d**
Bec. The energy of sublevel 4s is lower than that of 3d where sum of (n+l) of sublevel 4s is less than that of sublevel 3d

2- **The spin of a single electron in the same sublevel is in the same direction?**
To give the atom maximum stability and minimum energy

Sheet 3

1-Choose the correct answer for each of the following sentences:

1- The number of electrons that saturates the energy levels (M) and (N) are ……

a) 18 , 32 b) 8 , 18 c) 2 , 8 d) 8 , 32

2- When (n = 2) , the value of ℓ may equal ……

a) -1 b) 0 c) $-\frac{1}{2}$ d) 2

3- The possible values of n and ℓ for the same electron are …….. respectively.

a) 0 , 15 b) 3 , 3 c) 0 , 1 d) 2 , 1

4- The maximum number of energy levels in the heaviest known atom at the ground state is ……

a) 5 b) 6 c) 7 d) 8

5- The energy sublevels that from the third energy level are ……..

a) s b) s , p c) s , p , d d) s , p , d , f

6- The …… sublevel is not present in any atom.

a) 5d b) 1p c) 3p d) 2s

7- The value of the …….. quantum number for an electron that is located in the energy level L equals -1

a) principal b) subsidiary c) magnetic d) spin

8- The presence of 3 unpaired electrons in nitrogen atom $_7$N , in is stable state can be explained by ……

a) Pauli's exclusion principle b) Hund's rule

c) the uncertainty principle d) Aufbau principle

9- The atomic number of the element in which its electron configuration ends by $3d^2$ is ……

a) 18 b) 20 c) 22 d) 32

10- The quantum number which define the spin motion of the electron is.....

a. the principal quantum number

b. the subsidiary quantum number

c. the magnetic quantum number

d. the spin quantum number

8- Which configuration represents nitrogen according to Aufbau .....

d. $Is^2, 2s^1, 2p^4$. 9- **Heating** c. $Is^2, 2s^2, 2p_x^1\ 2p_y^1, 2p_z^1$ b. $Is^2, 2s^2, 2p^3$ a. 2 ,5

9-gases or vapours to high temperature under reduced pressure, they.....

b. emit light a. absorb energy

d. emit Alpha radiation c. emit gamma radiation

10 - When an electron absorb a quantum of energy it ..... a. Transfers to all higher energy levels

e. Transfers to the higher energy level corresponding to the absorbed energy

f. Transfers to lower energy level

g. Transfers to the lower energy level that corresponds to the absorbed quantum11-

11-The magnetic quantum number (m*l*) defines .....

a.The principal energy level.

b.The number of energy sublevels.

c.The number of orbitals and their shape.

d. The number of electrons in orbitals.

12- The no. of orbitals in the sublevel 3d equal **.....**

d. 7 c. 6 b. 4 a. 5

13- The no. of orbitals in the principal energy level (n) equals .....

d. (n-1) c. n^2 b. $3n^2$ a. $2n^2$

14- The maximum no. of electrons that occupy a given energy level (n) equals.....

d. $(2n)^2$ c. $2n^2$ b. n^2 a. 2n

15- The energy sublevel may be arranged according to their increasing energy in an ascending order as following:

b. 3s< 4p < 3d < 4f a. 3s< 3p<4d< 4s

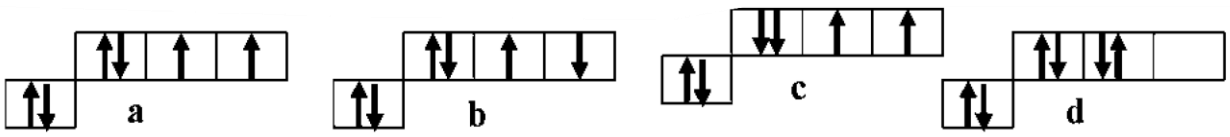
d. 3s<3p <4s<3d c. 3s< 3p < 3d < 4s

16- The orbitals of the same energy sublevel are .....

a. different in energy b. equal in energy

c. different in shape d. (a and c) correct

17 - One of the following diagrams shows the correct distribution of electrons in the last energy level of oxygen atom .



18-Which of the following quantum numbers for an electron include wring giving reason

a) n = 3 , ℓ = 2 , mℓ = -1, $m_s = +\frac{1}{2}$

b) n = 4 , ℓ = 3 , mℓ = -2 , $m_s = +\frac{1}{2}$

c) n = 1, ℓ = 1 , mℓ = 1, $m_s = -\frac{1}{2}$

<u>2-Give reason for each of the following:</u>

The principal quantum number is an integer number.

……
……
……

The third energy level is saturated with 18 electrons whereas the first level is saturated with 2 electrons.

……
……
……

5- Electron prefers pairing with another one in the same sublevel rather than entering the higher energy sublevel.

……
……
……

6- The 4s sublevel is filled with electrons before 3d sublevel.

……
……
……

4- The spin of single electrons in the orbitals of the same sublevel should be in the same direction before pairing.

……
……
……

8- The electron configuration of helium atom is $1s^1$, $2s^1$.

……
……
……

9-No electron pairing takes place until each orbital contains one electron.

……
……
……

10- The sublevel p takes up to 6 electrons whereas sublevel d takes 10 electrons.

……
……
……

3-Write the scientific expression for each statement of the following:

1) No electron pairing takes place in a given sublevels until each orbit contains one electron
2) Electrons occupied the energy sublevels in order of increasing their energy, the lowest sublevels are filled first
3) It is impossible for two electrons in the same atom to have the same four quantum numbers.
4) It represents the number of orbitals within a certain energy sublevel and their direction in space.
5) It indicates the number of energy sub levels in each principle level.
6) Bohr used it in hydrogen atom it is used to define Orders of principle energy levels
7) They are numbers define the volume of space (orbital) where there is maximum probability of finding electrons. Besides, they define the energy, shape and direction of orbitals.

4-Miscellaneous question:

1- What is the number of unpaired electron present in each atom of the following :

1) $_{7}N$

2) $_{16}S$

3) $_{26}Fe$

4) $_{11}Na^{+}$

5)$_{35}$ Br

6)$_{30}Zn$,

7) $_{18}Ar$,

8)$_{20}Ca$

9) $_{11}Na$

QUESTION 5:-

Write the possible values (ℓ) , (ml) for the electron its principle quantum number,(n = 2).

QUESTION 6:-

What are possible values of (ℓ) when (n=3) ?

QUESTION 7:-

What is meant by

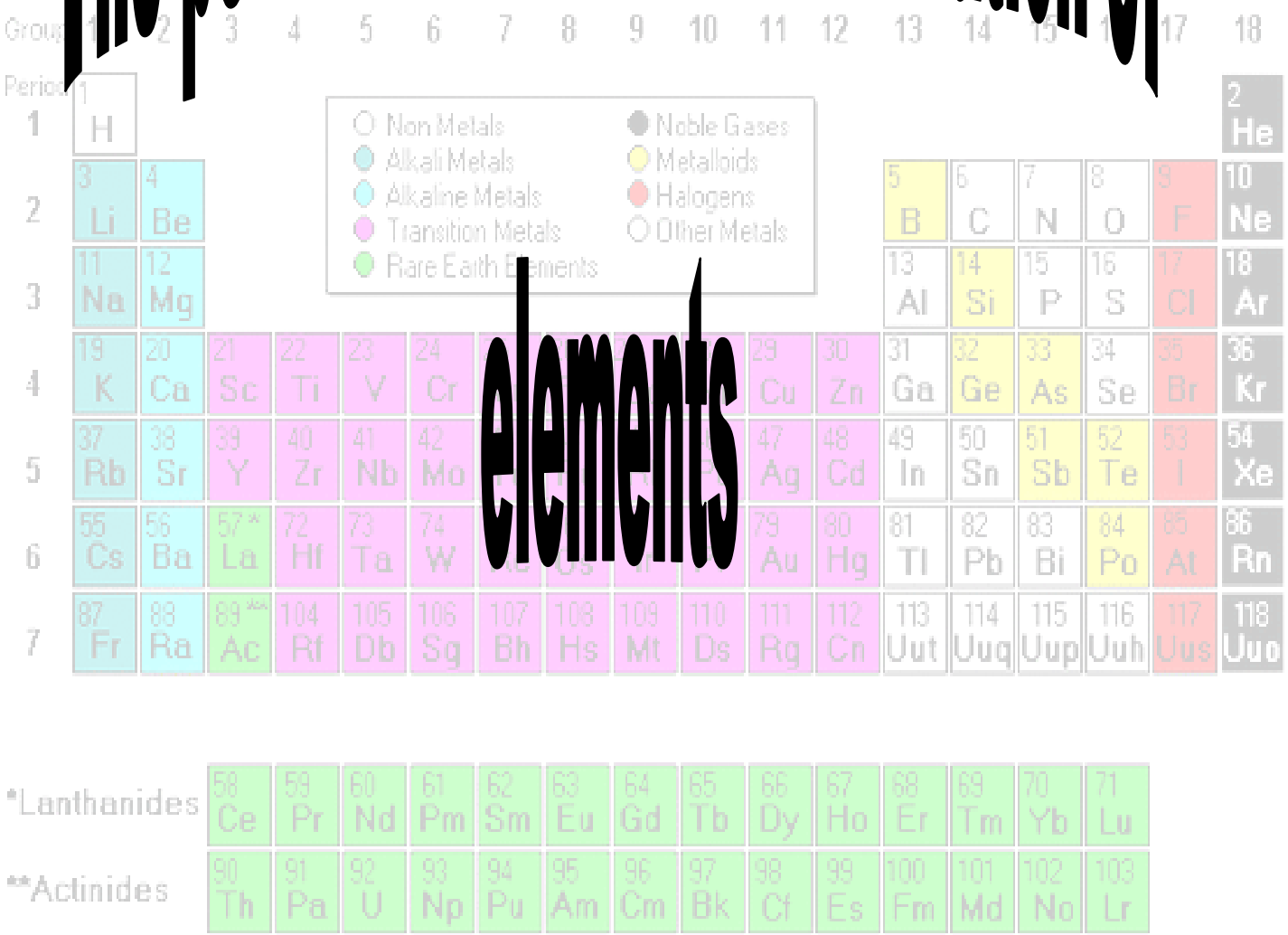
c. The building up principle

d. Hund's rule

Pauli exclusion principle

Chapter TWO

The periodic Table and classification of elements



Classification of elements and the long periodic table

The long form periodic table:

It depends on the building up principle (Auf - bau)

The elements are arranged ascendingly in the long form periodic table:

According to

1- their atomic numbers

The atomic number of each element increases than the element preceds it in the same period by whole one

2- the way to fill their atomic energy sublevels by electrons

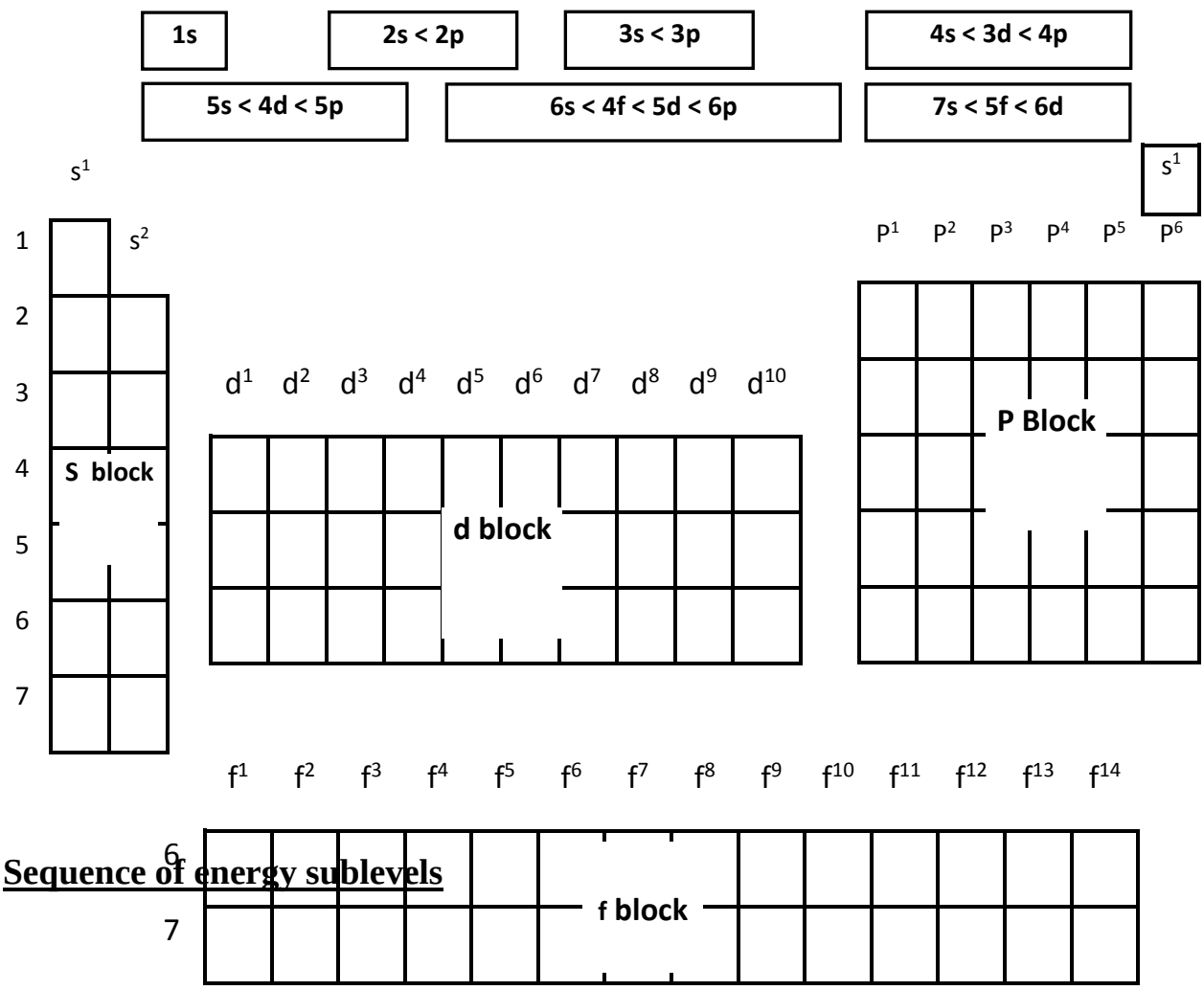
The periodic table is divided into four main blocks

1- S – block elements

2- P – block elements

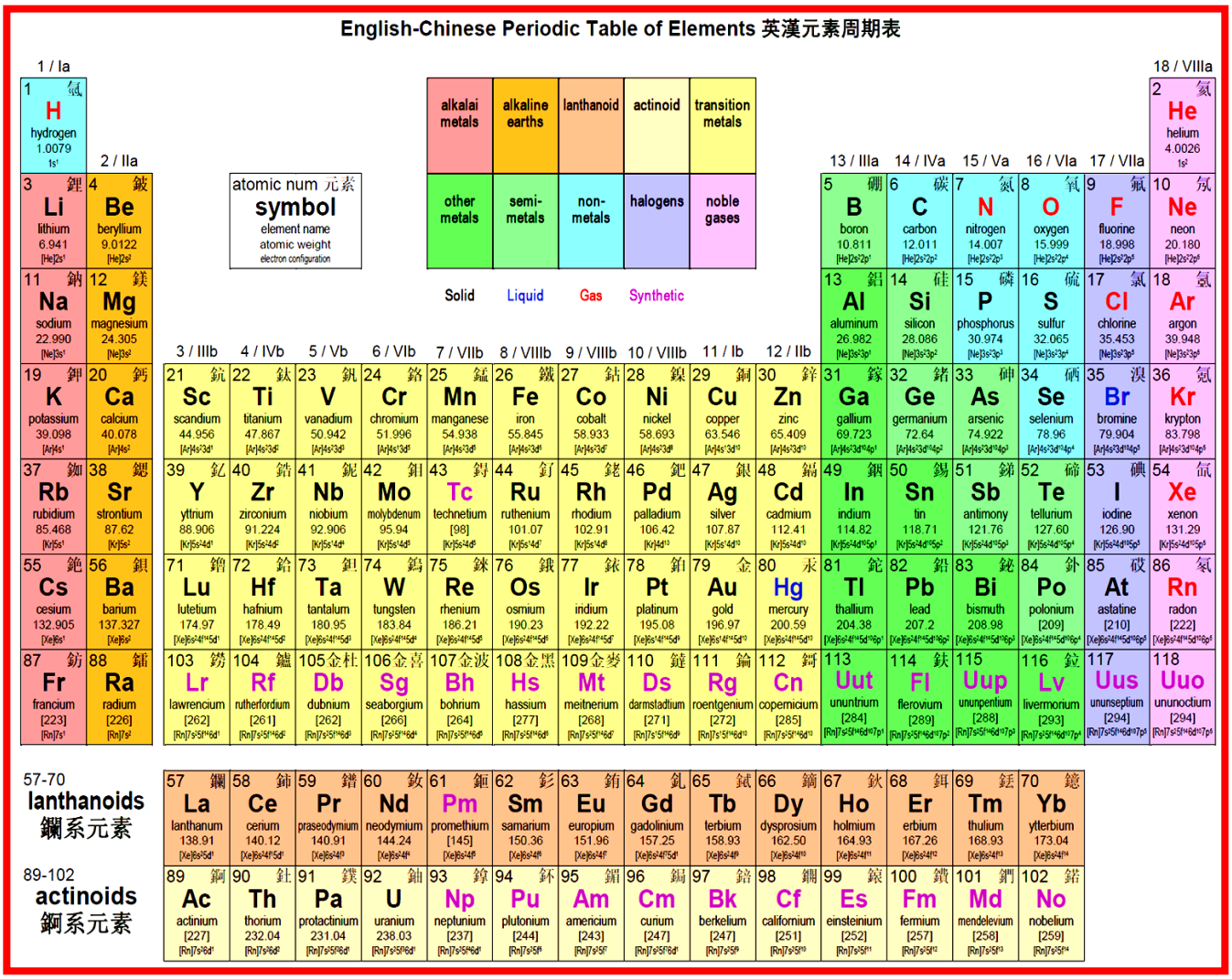
3- d – block elements

4- f – block elements



Sequence of energy sublevels

1S < 2S < 2P < 3S < 3P < 4S < 3d < 4P < 5S < 4d < 5P < 6S < 4f < 5d < 6P < 7S < 5f < 6d



Notes:

1-each period begins by filling a new energy level

2-the elements of the same vertical group are

a-identical in the electronic composition of their outer most level

b-different in the principal quantum no. (n)

3-the modern periodic table consists of 118 elements distributed in seven periods

Period	**1st**	**2nd**	**3rd**	**4th**	**5th**	**6th**	**7th**
No.of elements	2	8	8	18	18	32	32

Description of long periodic table:

S – Block

1- Placed in the left hand block of the table
2- It contains the elements whose outermost electrons occupy the s-sublevel
3- Consists of 2 groups (1A (ns1),2A (ns2),)

P – Block

1- Placed in the right hand block of the table
2- It contains the elements whose outermost electrons occupy the P-sublevel
3- Consists of 6 groups (3A (ns^3),4A (ns^4), 5A (ns^5) ,6A (ns^6), 7A (ns^7) ,0 (ns^8))

……

d – Block (they occupy the middle block of the table)

1- They occupy the middle block of the table
2- Contains the elements with the outermost electrons occupying the d- sublevel.
3- It contains ten vertical columns, seven of which belong to B-groups and the other three of them belong to group (VIII). Since the d-sublevel can take up to ten electrons
4- The d-block elements are known as transition elements and are themselves classified according to the number of the outer energy level and according to the period, giving three series which are :

1. The first transition series :

It includes the elements in which the sublevel 3 d is filled successively.

They are placed in the fourth period and consist of the elements from scandium to zinc.

2. The second transition series :

It includes the elements in which the sublevel 4d is filled successively.

They are placed in the fifth period and consist of the elements from yttrium to cadmium.

3. The third transition series :

It includes the elements in which the sublevel 5 d is filled successively.

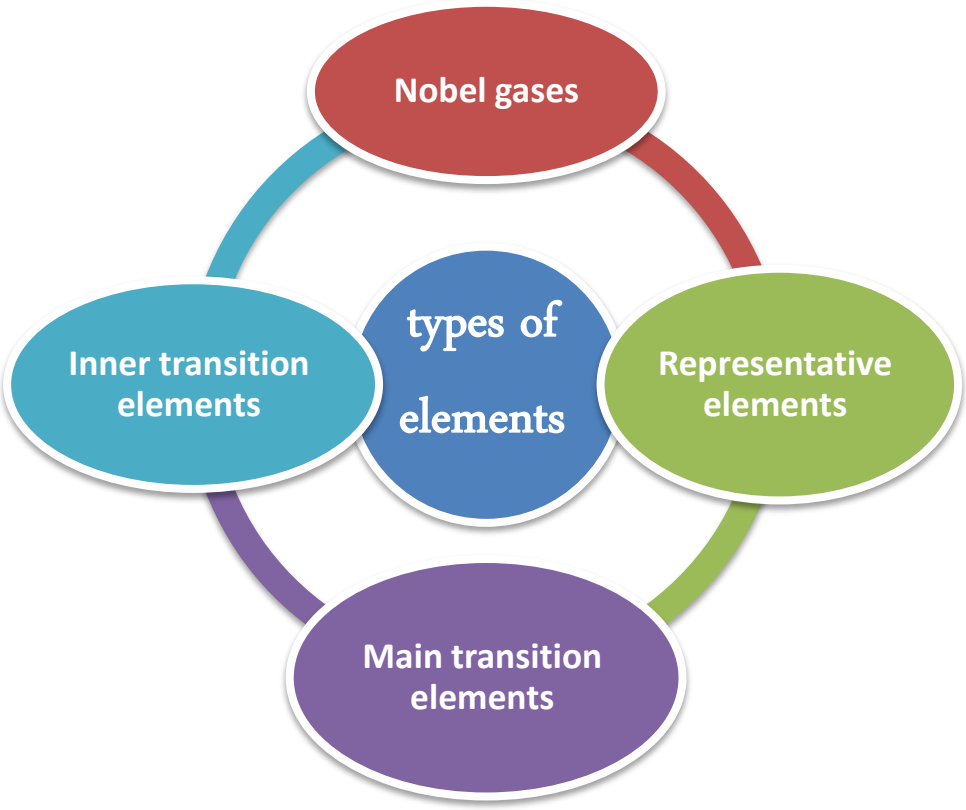
They are placed in the sixth period and consist of the elements from lanthanum to mercury.

d) f -Block elements :

1- This includes the elements in which the f-sublevel is filled successively, Sublevel f

2- Can take up to 14 electrons. The f-block includes two series i.e. the lanthanide and the actinide series.

The lanthanide series	The actinide series
• In this series the sublevel 4f is filled successively • it consists of 14 elements. • the outermost energy level for all these elements are $6s^2$, so they are quite similar in behavior. • they are very difficult to be separated and that is why they are known as rare earths. • They occupy the sixth period	• In this series the sublevel 5 f is filled successively • It includes 14 elements. • All the actinides are radioactive elements and their nuclei are unstable. • The f-block elements are known as the inner transition elements. • They are usually separated from the table (placed below it). So that it is not too wide. • They occupy the seventh period



Element	Definition
Nobel gases	They are the elements of the last column of the P-block (group zero)((18))their electronic structure is nP^6 except Helium is $1S^2$ and they are very stable elements
Representative elements	They are elements of S and P – blocks except that of group zero ,their energy levels are completely filled with electrons except the highest level and they tend to reach the completed configuration nS2, nP6 by gaining , losing or sharing electrons
Main transition elements	They are the elements of the d- block where the d-sublevels of these elements are successively filled They are characterized by having energy levels completely filled by electrons except the two outermost levels
Inner transition elements	They are the elements of the f- block where the f-sublevels of these elements are successively filled They are characterized by having energy levels completely filled by electrons except the three outermost levels

The abnormality of electron configuration for some elements of the periodic table:

Default elect. Config. Actual elect. Config.

$_{24}Cr$: [10Ne],$4S^2$, $3d^4$ ⟶ [10Ne],$4S^1$, $3d^5$

$_{29}Cu$: [18 Ar].$4S^2$, $3d^9$ ⟶ [10Ne],$4S^1$, $3d^{10}$

Interpretation

The atom becomes more stable when the d-sublevel being half-filled (d^5) as chromium or full-filled (d^{10}) as copper

The main transition elements	The inner transition elements
The elements of d-block	The elements of f-block
the d-sublevels of these elements are successively filled	the f-sublevels of these elements are successively filled
Contains three series	Contains two series
They appear in the fourth, fifth and sixth periods	They appear in the sixth and seventh periods
They are characterized by the completely filled energy levels with electrons except the outer two energy levels	They are characterized by the completely filled energy levels with electrons except the outer three energy levels
e.g : scandium – iron	e.g :cerium – thorium

The first transition series	The second transition series	The third transition series
Includes the elements in which the sublevel 3d is filled successively	Includes the elements in which the sublevel 4d is filled successively	Includes the elements in which the sublevel 5d is filled successively
Placed in the fourth period	Placed in the fifth period	Placed in the sixth period

How can you find the location and the type of element in the periodic table?
Write the electronic configuration of element in quantum levels.

2- Number of period = the maximum value of principle energy level (quantum number).

3- The no. of group = the no. of electrons in the outer main level

Element	Elec.config.	Block	Type	Period	Group
$_{12}$ **Mg**					
$_{23}$**V**					
$_{26}$**Fe**					
$_{29}$**Cu**					
$_{30}$**Zn**					
$_{17}$**Cl**					

Sheet 4

QUESTION 1:

Write the scientific expression for each statement of the following:

1- This includes the elements in which the f-sublevel is filled successively, Sublevel f

2-Placed in the left hand block of the table It contains the elements whose outermost electrons occupy the s-sublevel

3-Placed in the right hand block of the table It contains the elements whose outermost electrons occupy the P-sublevel

4- They occupy the middle block of the table Contains the elements with the outermost electrons occupying the d- sublevel.

5- It includes the elements in which the sublevel 3 d is filled successively They are placed in the fourth period

6- It includes the elements in which the sublevel 4d is filled successively

7- It includes the elements in which the sublevel 5 d is filled successively.

8- This includes the elements in which the f-sublevel is filled successively, Sublevel f

9- They are the elements of the last column of the P-block (group zero)((18))their electronic structure is nP^6 except Helium is $1S^2$ and they are very stable elements

10- They are elements of S and P – blocks except that of group zero ,their energy levels are completely filled with electrons except the highest level and they tend to reach the completed configuration nS^2, nP6 by gaining , losing or sharing electrons

11- They are the elements of the d- block where the d-sublevels of these elements are successively filled

12- They are the elements of the f- block where the f-sublevels of these elements are successively filled

13- The maximum value of principle energy level (quantum number).

14- the no. of electrons in the outer main level

QUESTION 2:

Choose from column (B) the electronic configuration of the element in column (A) then define the type of element in column (c).

Element	Electronic figuration	Type of element
1-Radon Rn_{86}	a. $7s^1$,	I. inner transition (actinide)
2- Cesium Cs_{55}	b. $6s^2$, $5d^6$	II. from 2nd transition series
3- Bromine Br_{35}	c. $6s^2$,$6p^6$	III. noble element
4- Vanadium V_{23}	d. $4s^2$,$3d^3$	IV.3^{rd} transition series
5-Molybdenum Mo_{42}	e. $6s^2$, $5d^1$, $4f^7$	V. inner transition (lanthanide)
6-Osmium Os_{76}	f. $4s^2$,$4p^5$	VI. representative from s- block
7-GadoliumGd_{64}	g. $5s^2$,$4d^4$	VII. first transition series
	i. $6s^1$	VIII. representative from p-block

QUESTION 3:

What is meant by each of the following:

1- Atomic number

2- Reduction

3- Representative element

4- Nobel element

5- Transition element

6- Inner transition element

QUESTION 4:

Choose the correct answer for each of the following sentences:

1- The long form periodic table consists of …… horizontal periods.

a) 7 b) 8 c) 10 d) 18

2- The following numbers are the atomic number of four elements, all are in the same period except with the atomic number of ……..

a) 3 b) 9 c) 10 d) 12

3- …….. elements exist in the same period.

a) $_{11}Na,_{10}Ne$ b) $_{11}Na,_{29}Cu$ c) $_{11}Na,_{3}Li$ d) $_{11}Na,_{17}Cl$

4- ……... elements are similar in the electron configuration.

a) Ne, Mg b) Na^+ , Ne c) Na, Mg d) H, Li

5- The elements with atomic number 4 and ……. have the same properties.

a) 8 b) 12 c) 19 d) 21

6- …….. is considered from transition elements.

a) $_{55}Cs$ b) $_{24}Cr$ c) $_{20}Ca$ d) $_{17}Cl$

7- The element with electron configuration : [Xe], $6s^2$, $4f^{14}$, $5d^3$ is considered from …….

a) the first transition series b) the third transition series

c) the lanthanides d) the representative elements

8- The number of actinides is …….. elements.

a) 4 b) 8 c) 14 d) 18

9- All the elements of the …… are radioactive elements and their nuclei are unstable.

a) lanthanide series b) last column of p-block

c) actinide series d) second transition series

QUESTION 5:*Give reason for each of the following:*

1- The properties of sodium ($_{11}Na$) and potassium ($_{19}K$) elements are similar.

……

……

……

2-Lanthanides are known as rare Earth's elements.

……
……
……

3-Lanthanides are quite similar in their chemical properties.

……
……
……

4- The abnormality of electron configuration chromium $_{24}Cr$ and copper $_{29}Cu$.

……
……
……

Trends and periodicity of properties in the periodic table

The atomic radius

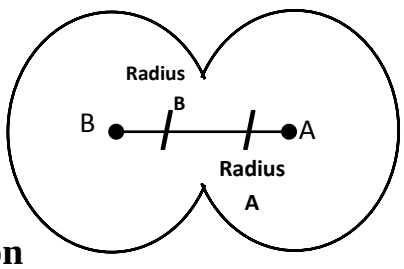
We cannot determine the atomic radius because

Electron has a wave motion so it is impossible to determine exactly the location of an electron around the nucleus.

The atomic radius:

It is half the distance between centers of two similar atoms in a diatomic molecule.

The bond length:



It is the distance between the nuclei of two bonded atoms.

There are many methods to measure the bond length such as :

1- X – ray. **2- Electron diffraction**.

.....

Examples

1- The bond length in the chloride molecule Cl – Cl is 1.98 Å and the length between carbon and chloride atoms C – Cl is 1.76 Å . Calculate the atomic radius of carbon.

solution

The atomic radius of chlorine = $\frac{1.98}{2}$ = 0.99 Å

The atomic radius of carbon = 1.76 – 0.99 = 0.77 Å

.....

3- The bond length in the molecule of NH3 is 1.0 Å and the bond length in the molecule of H2 is 0.6Å. Calculate the bond length in nitrogen molecule (N2)?

Solution

Note:

In ionic crystals constructed of cations and anions (e.g the sodium chloride crystal) found in a crystalline form,

Ionic bond length

The distance between the centers of the nuclei of two bonded ions

Example:

Calculate the bond length of sodium chloride knowing that the radius of sodium ion = 0.95 A^0 and that of chloride ion = 1.81 A^0

……

……

……

Effective nuclear charge (Z_{eff})

The actual nuclear charge (+ve charge) which affects on an electron in an atom

Give reasons:

The effective nuclear charge always less than the nuclear charge?

Due to the screening effect of electrons of the inner energy levels on a part of the nuclear charge affecting the electrons

The graduation of the atomic radius property in the periodic table

In the horizontal periods

Atomic radius **decrease** in period by increasing atomic number

Due to increasing The effective nuclear charge (Z_{eff})

.....

In the vertical groups

Atomic radius **increases** in group by increasing of atomic number why ?

- Due to increasing the number of energy level**s**
- Increasing the no. of the filled energy levels
- Increasing the screening effect
- Increasing of the repulsive forces between electrons

.....

Notes

1- the atoms of the 1st group elements are the biggest atoms

2- the atoms of the 7th group elements (halogens) are the smallest atoms

3- the biggest atom in size ic cesium (Cs) while the smallest one is fluorine (F)

.....

Notes

1- The cation (+ve ion) radius is smaller than that of its atom ?

This is due to the increasing of positive charges of the protons which attract the valency electrons leading to a decrease in the cation radius.

2- The anion's radius is bigger than that of its atom?

This is due to the increase in the number of the negative charges in the shells therefore the repulsive force between the electrons increases so the shells move a part and this leads to increasing of the anionic radius than that of its atom.

Ionization potential (Ionization energy)

It is the amount of energy required to remove the most bound electron completely from an isolated gaseous atom

M + Energy $\longrightarrow$ $M^+ + e^-$ **ΔH = + ve KJ/Mole**

The atom losses electrons and converted into positive ions . It has a positive value (ΔH = + ve KJ/Mole) .

.....

The first ionization energy:

It is the energy required to remove one electron from neutral atom to form a cation (+ ve) with one positive charge.

$M \longrightarrow M^+ + e^-$ **ΔH = + ve KJ/Mole**

The second ionization energy:

It is the energy required to remove one electron from neutral atom to form a cation (+ ve) with one positive charge.

$M^+ \longrightarrow M^{+2} + e^-$ **ΔH = + ve KJ/Mole**

- **It is greater than the 1st ionization energy**

The first ionization energy	The second ionization energy

2- The first ionization energy of noble gas is very high ?

Due to the stability of their electronic configuration because it is difficult to remove an electron from completely filled shell.

3- The ionization energy of element of group (5A)

(N_7 – P_{15} – As_{33} – Sb_{51} – Bi_{83}) is much greater than any element have the same period because the outer most energy sublevel (P) has three electrons and it is half filled with electrons (nP) and this gives the atom of the element some extra stability so the ionization energy is greater .

4- The ionization energy of sodium is much smaller than that of chlorine?

Because the atomic size of chlorine is smaller than that of sodium so the attractive force of the nucleus on the valence electrons in the case of chlorine is more strongly and the electrons valence need a higher energy to be separated from the atom.

5- Ionization energy increases period ?

Because the positive nuclear charge gradually increases with the increase of the atomic number led to decrease the atomic radius and increase the attractive force of the nucleus on the valence electrons therefore the electrons needed m large high energy to remove (separated) from the atom .

6- Ionization energy decreases in group ?

Due to the increase in the atomic size and screen of the attraction force of the nucleus on the valence electrons therefore the electrons needed a smaller value of energy to separate from the atom.

7- The ionization energy of elements of group (2A) is much greater than any element have the same period ?

Because the outer most energy level (S) of the element of group (2A) is completely filled with electron (nS^2) and this gives the atom of the element some extra stability so the ionization energy is much greater .

Electron affinity

It is the amount of energy released when an extra electron is added to a neutral gaseous atom to form an anion (- ve ion)

$X + e^- \longrightarrow X^- + energy \qquad \Delta H = -ve\ KJ/Mole$

Give reasons:

In the horizontal periods electron affinity increases with the increase in atomic number ?

Due to the atomic radius (size) gradual decrease so it becomes easier for the nucleus to attract the new electron .

The electron affinity decreases in group?

Due to the increase of the atomic volume with increase atomic number and this leads to the decreasing of the attraction force of the nucleus on the valence electrons .

Exception cases:

Beryllium has a relatively high of electrons affinity due to the stability of its atom that has completely filled orbitals ($1S^2$, $2S^2$) ?

Because the outer most energy sublevel (nS) is completely filled with electron and it gives the atom some extra stability .

Elements of the fifth group (N_7, P_{15}) have a lower value of electron affinity

Because the outer most energy sublevel (nP) has three electrons and it is half filled with electrons it gives the atom some extra stability (N_7 : $1S^2$, $2S^2$, $2P^2$).

Noble gases have not (small) electron affinity

Because all energy sublevels are completely filled with electrons which gives the atoms great stability .

Electron affinity of Fluorine (F_9) is less than that of chlorine (Cl_{17}) ?

Because the atomic radius (size) of fluorine atom is smaller than that of chlorine atom and when fluorine atom gains electron it is affected by a great repulsion force bigger than that in chlorine atom and fluorine atom is very small size .

Electro negativity

The tendency of an atom to attract the electrons of chemical bond to itself.

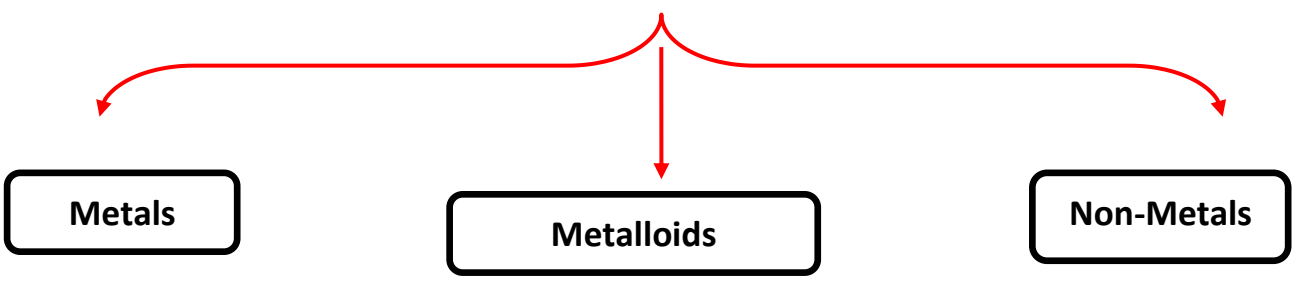
Notes

It is the average of the ionization potential and electron affinity

Electron affinity	Electro negativity
It is the amount of energy released when an extra electron is added to a neutral single atom to form an ion	It is tendency of an atom to attract the electrons of chemical bond to itself
It refers to the atom in its single state	It refers to the atoms which linked together in the molecule
Measured by Kj/mole	Measured by numbers till 4

Metallic and non-metallic property

Types of elements



Metals	Non metals
Valence shell has less than half its capacity of electrons (1 or 2 or 3)	Valence shell has more than half its capacity of electrons (5 or 6 or 7)
They ore called electropositive elements.	They are called electronegative elements
They have Relatively large atomic radius therefore the ionization energy and electron affinity and electro negativity have smell values	They have small atomic radius ionization energy and electro negativity and electron affinity have high value
They have a good electric conductor.	They have a bad electric conductor.
They lose electrons during the chemical reaction	They gain electrons during the chemical reaction

Metalloids

A group of elements that have a metallic appearance and most of the properties of nonmetals at the same time

Properties of metalloids

1. Electronegativity is intermediate between metals and nonmetals.
2. Their electrical conductivity is less than that of metals, but more than that of nonmetals.
3. They are used as semiconductors and are known as transistors.

.....

Notes

Metals are placed to the left of metalloids while nonmetals are placed to the right of metalloids.

Examples of metalloids:

Boron (B)- Silicon (Si) – Germenium (Ge) – Arsenic(As)- Antimony (Sb)- Tellurium (Te)

Metallic and non-metallic character trend in the periodic table as follows;

a) In periods:

As we move across the period from left to right we observe that the first group includes the elements of the highest metallic character.

Then this properly decreases gradually with the increase in the atomic number across the period past the metalloids. To the right of the metalloids begins the nonmetallic character. Group seven includes the elements of the highest nonmetallic character

……

b) In groups:

The metallic character increases with the increase in the atomic number in descending group. Consequently, we conclude that **the elements of strongest metallic** character are placed **at the bottom on the left hand side** of the table.

Thus **cesium i**s considered as the element which has **the highest metallic** character.

On the other hand, the elements which the highest nonmetallic character is found at the top of the right side of the table.

Thus **fluorine** is considered as the element which **has the highest nonmetallic character.**

Fluorine is the strongest non metal while cesium is the strongest metals

Acidic – basic properties

Acidic oxides	Basic oxides
They are non – metallic oxides such as CO_2 , SO_2 , SO_3 , P_2O_5	They are metallic oxides such as Na_2O , K_2O , MnO , CaO , BaO
They dissolve in water to form acids $Co_2 + H_2O \longrightarrow H_2CO_3$ $SO_2 + H_2O \longrightarrow H_2SO_3$ $SO_3 + H_2O \longrightarrow H_2SO_4$ $P_2O_5 + H_2O \longrightarrow 2H_3PO_4$	Some basic oxides dissolve in water to form alkalies and others are not $Na_2O + H_2O \longrightarrow 2NaOH$ $K_2O + H_2O \longrightarrow 2KOH$ $CaO + H_2O \longrightarrow Ca(OH)_2$ $MgO + H_2O \longrightarrow Mg(OH)_2$
They react with alkalis to form salt and water $CO_2 + 2NaOH \longrightarrow Na_2CO_3 + H_2O$ $SO_2 + 2NaOH \longrightarrow Na_2SO_4 + H_2O$	They react with acids to form salt and water $Na_2O + 2Hcl \longrightarrow H_2O + 2Nacl$ $MgO + H_2SO_4 \longrightarrow MgSO_4 + H_2O$
They do not reaction with acids	They do not react with alkalis

<u>There is a third type of oxides known as Amphoteric</u>:

like aluminium oxide AI_2O_3, zinc oxide ZnO, antimony oxide Sb_2O_3 and tin oxide SnO. These oxides react either as basic oxides or as acidic oxides.

$$ZnO_{(s)} + H_2SO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + H_2O_{(\ell)}$$

$$ZnO_{(s)} + 2NaOH_{(aq)} \longrightarrow Na_2ZnO_{2(aq)} + H_2O_{(\ell)}$$

Sodium zincate

a) In periods:

We observe the acidic character in oxides increases when the atomic number increases and the basic character decreases.

……

b) In groups:

If we consider the elements of the first group as example to explain the trends in the **basic property**.

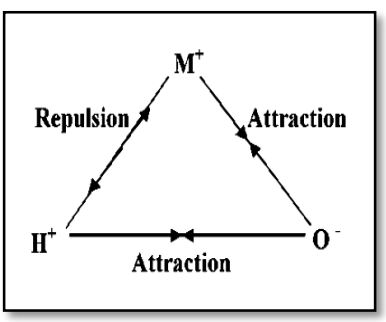
We find that **it increases** in **descending the groups** or with **increasing atomic number.**

This is due to the increases in the atomic size of the element, while the charge remains constant.

Considering acids and bases as hydroxy compounds, they may be represented by the general formula (MOH), (where M is the element atom), It may be ionized by either ways:

1- It may produce hydroxide ions and considered a base

$MOH_{(Base)} \longrightarrow M^+ + OH^-$



2- Or it may produce hydrogen ions and considered an acid

$MOH_{(Acid)} \longrightarrow MO^- + H^+$

Assume the three atoms of MOH are arranged in a triangle as follows :

a. If the attraction force between $M^+ O^-$ is bigger than that between H^+ , O^- the substance will ionize as an acid.

b. If the attraction force between H^+ and O^- is bigger than that between M^+ and O^-, the substance will ionize as a base.

c. If the attraction forces are equal, the substance will ionize as an acid or a base depending on the reaction medium this means that it reacts as base in the acidic medium and as an acid in the basic medium.

The attraction forces in the previous reactions depends on the volume and the charge of the M4 atom . In alkali metals like sodium, we observe that sodium atom has a big volume however, it has only one positive charge.

Accordingly the attraction between Na^+ and O^- is weaker than the attraction between O- and H+ so OH^- ion is produced i.e it is ionized as a base . However, if we move in the same period to the right, we observe that the nonmetal atoms as chlorine has a small volume and a big charge which increases its attraction to O- and ionized as an acid . The strength of oxygenated acids depends on the number of oxygen atoms which are not linked to hydrogen atoms.

If we represent the oxygenated acid by the formula [MOn(OH)m], where M is the element atom , we observe that the strong acid is that which has more number of non bonded oxygen atoms (On) with hydrogen.

Acid MOn(OH)m	No.of nonbinded Oxygen atoms with hydrogen	Strength of the acid
Orthosiliconic H_4SiO_4 $Si(OH)_4$	Zero	Weak
Orthophosphoric H_3PO $PO(OH)_3$	1	Moderate
Sulphuric H_2SO_4 $SO_2(OH)_2$	2	Strong
Perchloric $HC\ell O_4$ $C\ell O_3(OH)$	3	Very strong acid

Oxidation numbers

Oxidation number:

It is a number that refers to the electric charge (positive or negative) that the atom or ion would have in the compound, be it ionic or covalent"

Rules for assigning oxidation numbers

1-In ionic compounds:

The oxidation number of any atom is equal to its valency preceded by a positive sign in case of cations and with a negative one in case of anion. If the oxidation number is positive, this indicates that the number of electrons that the atom has lost to give this. If the oxidation number is negative this indicates that the number of the electrons that the atom has gained to give this.

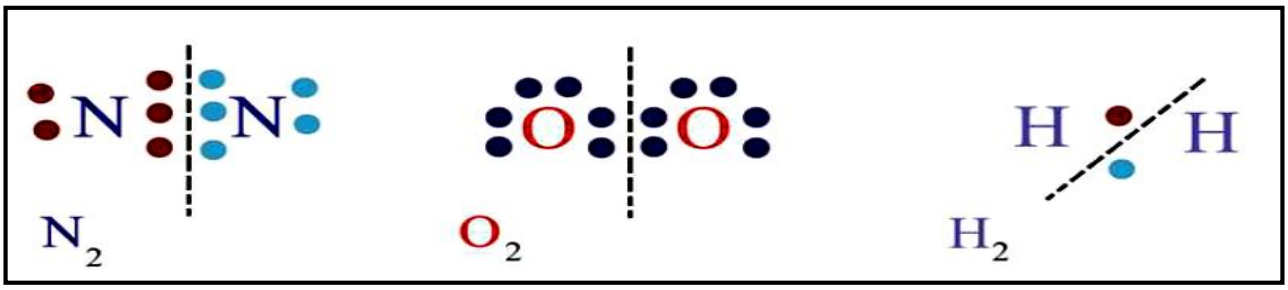
For example: K^+Br^-, $Na^+C\ell^-$, $Mg^{2+}O^{2-}$, $Ca^{2+}(CO3)^{2-}$, $Cu^{2+}(SO4)^{-2}$

2- In covalent compounds:

Since there are no negative or positive ions, we consider the charge carried on the atom explains the electronic shift in the chemical bond. The more electronegative atom carries a negative charge and the less electronegative one carries a positive charge. There are two cases in assigning the oxidation number in covalent compounds they are :

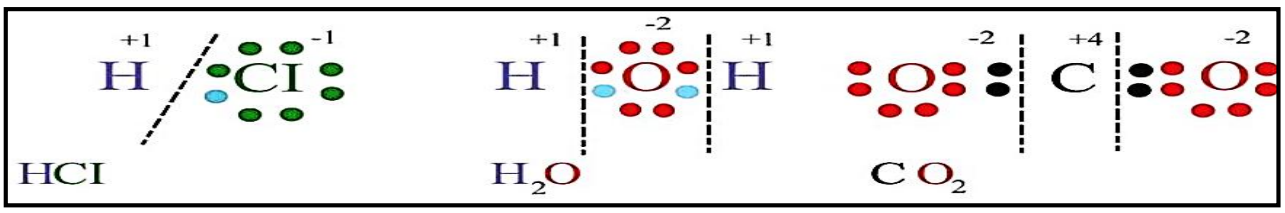
a. In molecules of similar atoms

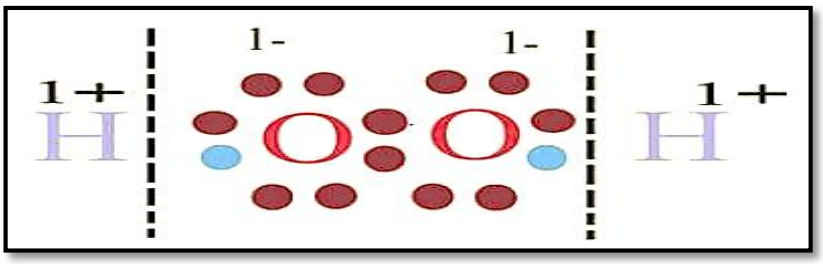
e.g. $C\ell_2$,O_3,P_4,S_8 . the electronic shift in the bonds between the atoms are equal, because the electronegativity of the atoms forming the molecule are similar, Accordingly, the oxidation number of any atom in this molecule is zero.



b) In diatomic molecules of different atoms in electronegativity molecules of similar atoms

, the shared electrons are assigned to the more negative atom.





It must be noticed that:

1- Oxygen

in most of its compounds has an oxidation number o **(-2)**

Except for few compounds such as

Peroxides (e.g.) **hydrogen peroxide** (oxygen water) H_2O_2 in which oxygen has an oxidation number of **(-1)**

2- Hydrogen

normally has an oxidation of **+1** in its compounds

Except in binary compounds with active metals (**metal hydrides**),

eg. **sodium hydride** Na H and calcium hydride CaH2 .

Hydrides

They are ionic compounds in which hydrogen is the negative ion.

Note:

If sodium hydride is melted and electrolyzed **Hydrogen** gas will evolve at **the anode** Bec. The oxidation number of Hydrogen in hydrides (-1)

In addition to these basic rules, the below also prove useful:

1. The algebraic sum of the oxidation number of all atoms in a neutral compound is zero.

2. The oxidation number is counted to one atom or ion only in the molecule.

3. The oxidation number of group IA elements is always (+1), group IIA elements is always (+2) and that for group IIIA elements is (+3).

Accordingly, on calculating the oxidation numbers of the atoms in a given compound we start by assigning the oxidation number of the elements of these groups, then we complete the other atoms.

Compound	Na_2O Sodium oxide	Na_2O_2 sodium peroxide	KO_2 Potassium super oxide	CaH_2 Calcium hydride	AlH_3 Aluminum hydride
Total of charges	+2 -2	+2 -2	+1 -1	+2 -2	+3 -3
The atom	Na_2 O	Na_2 O_2	K O_2	Ca H_2	Aℓ H_3
Oxdation no. of each atom	+1 -2	+1 -1	+1 -1/2	+2 -1	+3 -1

4. The algebraic sum of the oxidation number of all atoms in a polyatomic ion is equal to the charge of the ion.

Some common polyatomic ions are:

$(NO3)^{-1}$ nitrate, $(CO3)^{-2}$ carbonate, $(SO4)^{-2}$ sulphate, $(NH4)^{+1}$ ammonium ion.

The advantage of using oxidation number

is that they can help us to tell the type of chemical change occurring to an element during the chemical reaction. For example, in oxidation and reduction reactions,

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Oxidation

The process of losing electrons resulting in an increase of the positive charge;

Reduction

The process of gaining electrons resulting in a decrease of the positive charge.

By following the oxidation number in a chemical reaction, we can recognize the oxidation or reduction process.

Exercise:

potassium dichromate reacts with iron (II) chloride (ferrous chloride) according to the equation

$K2Cr2O7 + 6FeCℓ2 + 14HCℓ \longrightarrow 2KCℓ + CrCℓ3 + 6FeCℓ3 + 7H2O$

Explain the type of change (oxidation or reduction) that occurred to chromium and iron in this reaction.

Sheet 5

Choose the correct answer for each of the following sentences:

1- The biggest atom in size in a given period is that of …… group.

a) 1A b) 1B c) halogens d) 18

2- The biggest atom in size through the third period is …… atom.

a) $_{17}Cl$ b) $_{15}P$ c) $_{13}Al$ d) $_{11}Na$

3- The ionic radius of the metal is ……. its atomic radius.

a) bigger than b) smaller than c) equal to

4- The equation: $X+e^- \rightarrow X^-$ + energy , represents……

a) The electron affinity concept
b) a chemical process whose ▲H is positive
c) an endothermic process
d) the first ionization energy concept

5- Which of the following elements has the highest first ionization potential ? ……

a) $_8O$ b) $_7N$ c) $_6C$ d) $_5B$

6- The electron configuration for the atom which has the highest second ionization potential is ……..

a) $1s^2 , 2s^2 , 2p^6 , 3s^2$ b) $1s^2 , 2s^2 , 2p^5$

c) $1s^2 , 2s^2 , 2p^6$ d) $1s^2 , 2s^2 , 2p^6 , 3s^1$

7- The equation …….. represents the second ionization potential for the element X.

a) $X_{(g)} \longrightarrow X2^+_{(g)} + 2e^-$ b) $X^+_{(g)} \longrightarrow X2^+_{(g)} + e^-$

c) $X_{(g)} + 2e^- \longrightarrow X2^-_{(g)}$ d) $X^-_{(g)} + e^- \longrightarrow X2^-_{(g)}$

8- ……. has the higher electronegativity.

a) Lithium b) Fluorine c) Hydrogen d) Caesium

9- Three different elements A,B and C are in the same period and three successive group, if the element C is a noble gas , the symbol of A ion will be ……

a) A^+ b) A^{++} c) A^- d) A^-

10- …….. classified the elements into metals and nonmetals.

a) Bohr b) Mendeleev c) Berzelius d) Thomson

Write the scientific expression for each statement of the following:

1) It is half the distance between centers of two similar atoms in a diatomic molecule.
2) It is the distance between the nuclei of two bonded atoms
3) The actual nuclear charge (+ve charge) which affects on an electron in an atom
4) It is the amount of energy required to remove the most bound electron completely from an isolated gaseous atom
5) It is the energy required to remove one electron from neutral atom to form a cation (+ ve) with one positive charge.
6) It is the energy required to remove one electron from neutral atom to form a cation (+ ve) with one positive charge.
7) It is the amount of energy released when an extra electron is added to a neutral gaseous atom to form an anion (- ve ion)
8) The tendency of an atom to attract the electrons of chemical bond to itself.

Give reason for each of the following:

1- It is incorrect to define the atomic radius as the distance from the nucleus to the farthest electron.

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2- Atomic radius of chlorine atom ($_{17}Cl$) is smaller than that of sodium atom ($_{11}Na$).

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3- The atomic radii of the elements decrease as we move from left to right through the period.

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4- The bond length of $Fecl_3$ molecule is shorter than that of $FeCl_2$ molecule.

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5- The first ionization energy of noble gases is very high.

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6- The ionization potential of aluminum $_{13}Al$ is lower than that of magnesium $_{12}Mg$, although aluminum comes next magnesium in the same period.

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…….....

7- It is easy to oxidize iron (II) into iron (III) , but it is difficult to oxidize manganese (II) into manganese (III).

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…….....

8- The electron affinity of group (18) elements has no value.

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…….....

9- The electronegativity decreases as we go down the group.

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…….....

Problems:

1- Calculate the atomic radius of the sulphur atom providing that :

- The radius of hydrogen atom equals 0.3A° and the bond length of H_2S molecule equal 1.34A°.

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2- Calculate the atomic radius of the hydrogen atom providing that:

- The bond length (O – H) in water molecule = 0.96 A°.
- The bond length in oxygen molecule (O_2) = 1.32 A°.

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Sheet 6

QUESTION 1:

Choose the correct answer for each of the following sentences:

1- Metalloids are characterized by …….

a) Valence shell ,filled with less than that half of its capacity.

b) high electronegativity.

c) electric conductivity more than that of the metals.

d) the appearance of metals and properties of nonmetals.

2- Most of metallic oxides are …… oxides.

a) per b) amphoteric c) basic d) acidic

3- All the following oxides are amphoteric oxides except ……

a) Sb_2O_3 b) ZnO c) Al_2O_3 d) As_2O_3

4- Orthophosphoric acid H_3PO_4 …….

a) is trihydroxyl acid. b) is very strong acid.

c) reacts with nonmetal oxides forming salt and water.

d) its phosphorus atom combines with three oxygen atoms unlinked with hydrogen.

5- The oxidation number of chlorine in chlorine molecule is equal……..

a) +7 b) +1 c) -1 d) zero

6- The oxidation number of hydrogen in ……. Is equal -1

a) CaH_2 b) H_2O c) H_2O_2 d) HCl

7- The reaction : $2FeSO_4 \xrightarrow{\Delta} Fe_2O_3 + SO_2 + SO_3$, represents …….

a) reduction of iron. b) oxidation of sulphur.

c) reduction of oxygen. d) oxidation of iron and reduction of sulpher.

8-Nonmetals are characterized by .....

a. large ionization energy b. electropositive elements

c. small electron affinity d. large atomic radius

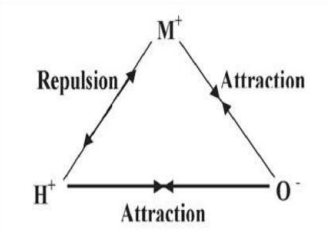
9-Electronegativity increases across the periods by .....

a. increasing atomic radius b. decreasing the atomic number

c. decreasing atomic radius d. (a and b) correct.

10-In the shown diagram, if the attraction between O^- and M^+ is greater than that between O^- and H^+, the substance is ionized as a .....

a. base b. acid c. acid and base d. unionized



11.In the previous diagram ,If M^+ is sodium

a- o^- attracts more to hydrogen ion.

b- o^- attracts more to sodium ion.

c- the substance ionized as an acid d- bound stronger between O and sodium

12. One of the following reactions does not represent an oxidation reduction reaction , it is no .....

a. $2P + 5HC\ell O + 3H_2O \longrightarrow 2H_3PO_4 + 5HC\ell$

b. $Zn + 2HC\ell \longrightarrow ZnC\ell_2 + H_2$

c. $Mg + CuSO_4 \longrightarrow MgSO_4 + Cu$

d. $NaOH + HNO_3 \longrightarrow NaNO_3 + H_2O$

13. One of the following reaction represents an oxidation reduction reaction , it is no .....

a. $CuO + H_2SO_4 \longrightarrow CuSO_4 + H_2O$

B. $CaCO_3 + 2HC\ell \longrightarrow CaC\ell_2 + H_2O + CO_2$

c. $Cr_2O^{2-}_7 + 3H_2S + 8H^+ \longrightarrow 2Cr^{3+} + 3\,S + 7H_2O$

d. $NaC\ell + AgNO_3 \longrightarrow Ag\,C\ell + NaNO_3$

QUESTION 2:

Give reason for each of the following:

1- The metals are good electric conductors whereas the nonmetals are bad ones.

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2- Cesium has the highest metallic character in group (1A).

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3- Fluorine is considered the most active nonmetal.

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4- Zinc oxide reacts with NaOH , although it is a metal oxide.

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5- $ClO_3(OH)$ acid is stronger than HCl acid.

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6- On electrolysis of molten sodium hydride, hydrogen evolves at anode but it evolves at cathode on electrolysis of acidified water.

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7- The oxidation number of nitrogen in the oxygenated compounds is positive, whereas it is negative in the hydrogen compounds.

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8- The oxidation number of oxygen in oxygen fluoride is positive.

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<u>Write the scientific expression for each statement of the following:</u>

1- The elements whose valence energy level is nearly half filled with electrons.

2 - The elements which are used in manufacturing of transistors as they are semiconductors.

3- The metallic oxides which dissolve in water.

4- The oxides that react either as basic oxides or as acidic oxides.

5- The compounds in which the oxidation number of hydrogen is (-1).

QUESTION 8:

Calculate the oxidation number of the mentioned element in following compounds :

a. oxygen in : O_2 - O_3 - Li_2O - Na_2O_2 - KO_2 - OF_2.

b. chlorine in : Na Cℓ - Na $C\ell O_4$ - Na $C\ell O_3$ - Na $C\ell O_2$ - Na CℓO .

c. nitrogen in :NH_3 - N_2 - N_2O - NO - HNO_2 - NO_2 - HNO_3.

d. sulphur in : $Na_2S_2O_3$ - K_2S - SO_2 - $NaHSO_3$ - H_2SO_4.

e. manganese in : MnO_2 - $KMnO_4$ - $MnC\ell_2$ - $NaMnO_4$.

QUESTION 9:

By examining changes in oxidation number in the following reactions, decide whether an oxidation or reduction process has taken place.

a) CO ⟶ CO_2

b) $Cr_2O_7^{2-}$ ⟶ Cr_2O_3

c) O_2 ⟶ O_3

d) NO_2 ⟶ N_2O_4

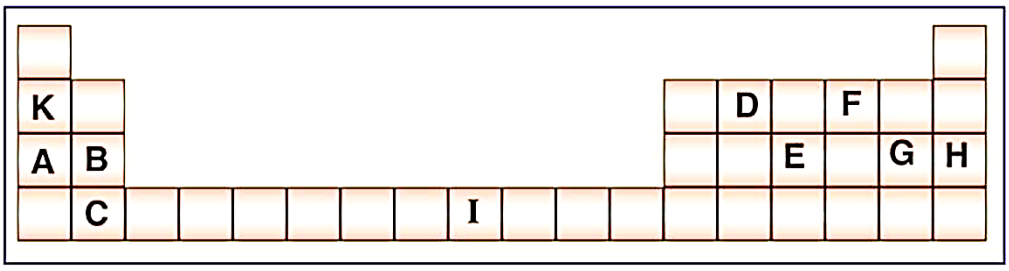
e) MnO^{4-} ⟶ MnO_2

f) $C\ell O^-$ ⟶ ClO^{3-}

g) $FeCl_3$ ⟶ $FeCl_2$

QUESTION 10:

The following diagram represents the first four periods of the long form periodic table :



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QUESTION 11:

Show by balanced symbolic equation that:

a. Sodium oxide is from basic oxide.

b. Sulphur trioxide is from acidic oxide.

c. Zinc oxide is from amphoteric oxide.